

DOT POINT

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IB CHEMISTRY AHL



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Science Press

© Science Press 2010
First published 2010

Science Press
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What the book includes

This book provides questions and answers for each dot point in the IB Chemistry Additional Higher Level (AHL) syllabus from the International Baccalaureate Diploma Programme for Chemistry:

- Atomic Structure
- Periodicity
- Bonding
- Energetics
- Kinetics
- Equilibrium
- Acids and Bases
- Oxidation and Reduction
- Organic Chemistry

Format of the book

The book has been formatted in the following way:

1.1 Subtopic from syllabus.

1.1.1 Assessment statement from syllabus.

1.1.1.1 First question for this assessment statement.

1.1.1.2 Second question for this assessment statement.

The number of lines provided for each answer gives an indication of how many marks the question might be worth in an examination. As a rough rule, every two lines of answer might be worth 1 mark.

How to use the book

Completing all questions will provide you with a summary of all the work you need to know from the syllabus. You may have done work in addition to this with your teacher as extension work. Obviously this is not covered, but you may need to know this additional work for your school exams.

When working through the questions, write the answers you have to look up in a different colour to those you know without having to research the work. This will provide you with a quick reference for work needing further revision.

Command Terms and Verbs to Watch

account, account for State reasons for, report on, give an account of, narrate a series of events or transactions.

analyse Interpret data to reach conclusions.

annotate Add brief notes to a diagram or graph.

apply Use an idea, equation, principle, theory or law in a new situation.

assess Make a judgement of value, quality, outcomes, results or size.

calculate Find a numerical answer showing the relevant stages in the working (unless instructed not to do so).

clarify Make clear or plain.

classify Arrange into classes, groups or categories.

comment Give a judgement based on a given statement or result of a calculation.

compare Give an account of similarities and differences between two (or more) items, referring to both (all) of them throughout.

construct Represent or develop in graphical form.

contrast Show how things are different or opposite.

deduce Reach a conclusion from the information given.

define Give the precise meaning of a word, phrase or physical quantity.

demonstrate Show by example.

derive Manipulate a mathematical relationship(s) to give a new equation or relationship.

describe Give a detailed account.

design Produce a plan, simulation or model.

determine Find the only possible answer.

discuss Give an account including, where possible, a range of arguments for and against the relative importance of various factors, or comparisons of alternative hypotheses.

distinguish Give differences between two or more different items.

draw Represent by means of pencil lines.

estimate Find an approximate value for an unknown quantity.

evaluate Assess the implications and limitations.

examine Inquire into.

explain Give a detailed account of causes, reasons or mechanisms.

extract Choose relevant and/or appropriate details.

extrapolate Infer from what is known.

identify Find an answer from a given number of possibilities.

justify Support an argument or conclusion.

label Add labels to a diagram.

list Give a sequence of names or other brief answers with no explanation.

measure Find a value for a quantity.

outline Give a brief account or summary.

predict Give an expected result.

propose Put forward a point of view, idea, argument, suggestion etc for consideration or action.

recall Present remembered ideas, facts or experiences.

show Give the steps in a calculation or derivation.

sketch Represent by means of a graph showing a line and labelled but unscaled axes but with important features (for example, intercept) clearly indicated.

solve Obtain an answer using algebraic and/or numerical methods.

state Give a specific name, value or other brief answer without explanation or calculation.

suggest Propose a hypothesis or other possible answer.

summarise Express concisely the relevant details.

synthesise Put together various elements to make a whole.



Atomic Structure

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12.1.1	Ionisation energies and energy levels.	12.1.6	Using the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations.
12.1.2	Successive ionisation energies and electron configuration.		
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AHL 12

Atomic Structure



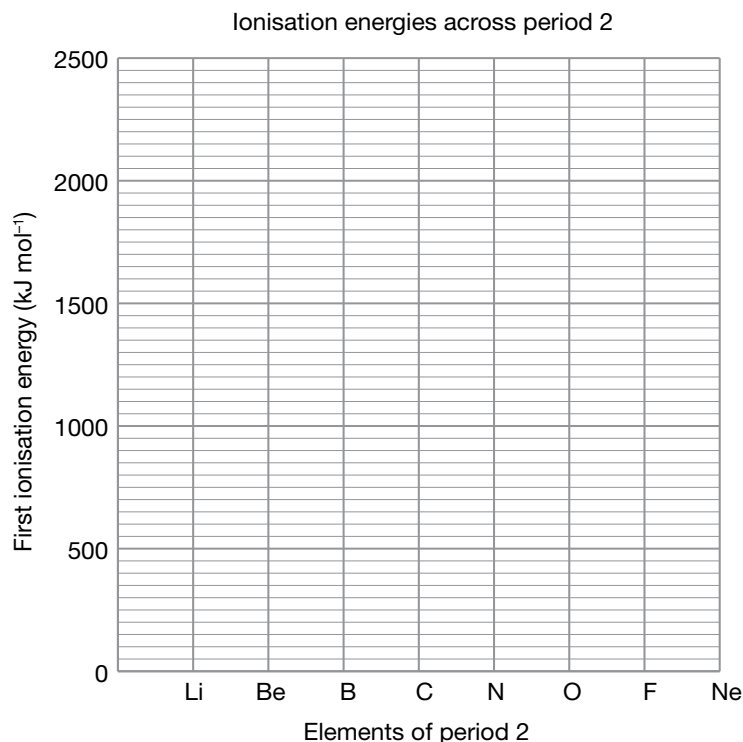
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12.1 Electron configuration. © IBO 2007

12.1.1 Explain how evidence from first ionisation energies across periods accounts for the existence of main energy levels and sublevels in atoms. © IBO 2007

12.1.1.1 Ionisation energy refers to the minimum amount of energy required to remove a mole of electrons from a mole of gaseous atoms in order to form a mole of gaseous ions.

- (a) Use information from the data tables in the appendices at the back of this book to graph the first ionisation energies of the elements across period 2 of the periodic table.



- (b) Based on this information, describe the trend in ionisation energy across period 2.

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- (c) Describe three factors that affect the ionisation energy of atoms.

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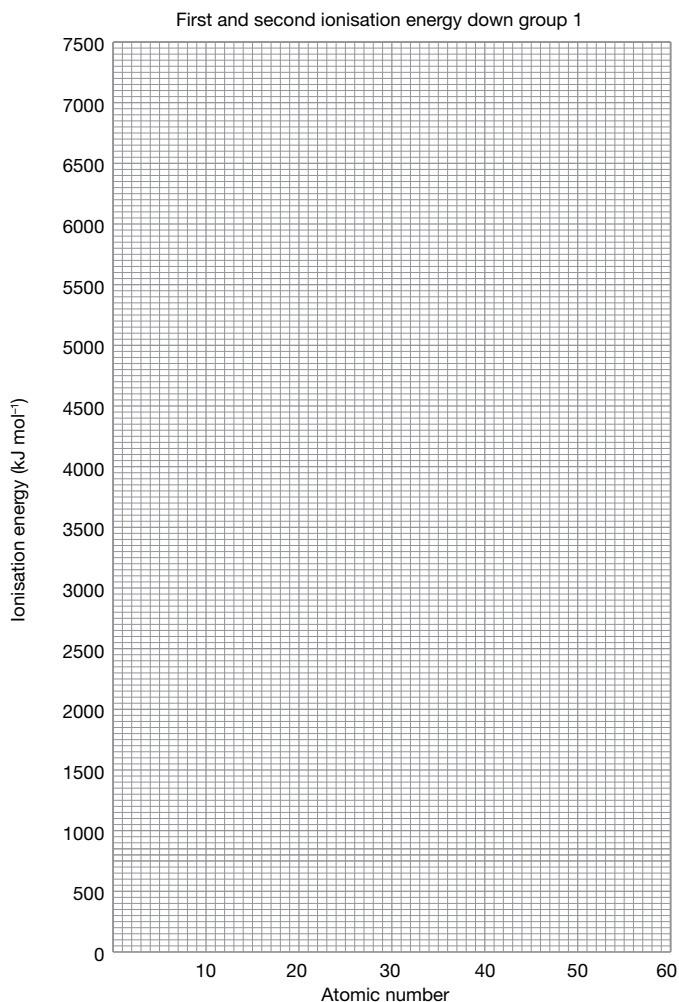
- (d) Describe and account for any variation from the trend of increasing ionisation across period 2.

12.1.2 Explain how successive ionisation energy data is related to the electron configuration of an atom. © IBO 2007

Element	Atomic number	First ionisation energy (kJ mol^{-1})	Second ionisation energy (kJ mol^{-1})
Lithium	3	520	7300
Sodium	11	496	4570
Potassium	19	419	3080
Rubidium	37	403	2660
Caesium	55	376	

12.1.2.1 The table provides information on the first and second ionisation energy values for elements in group 1 of the periodic table.

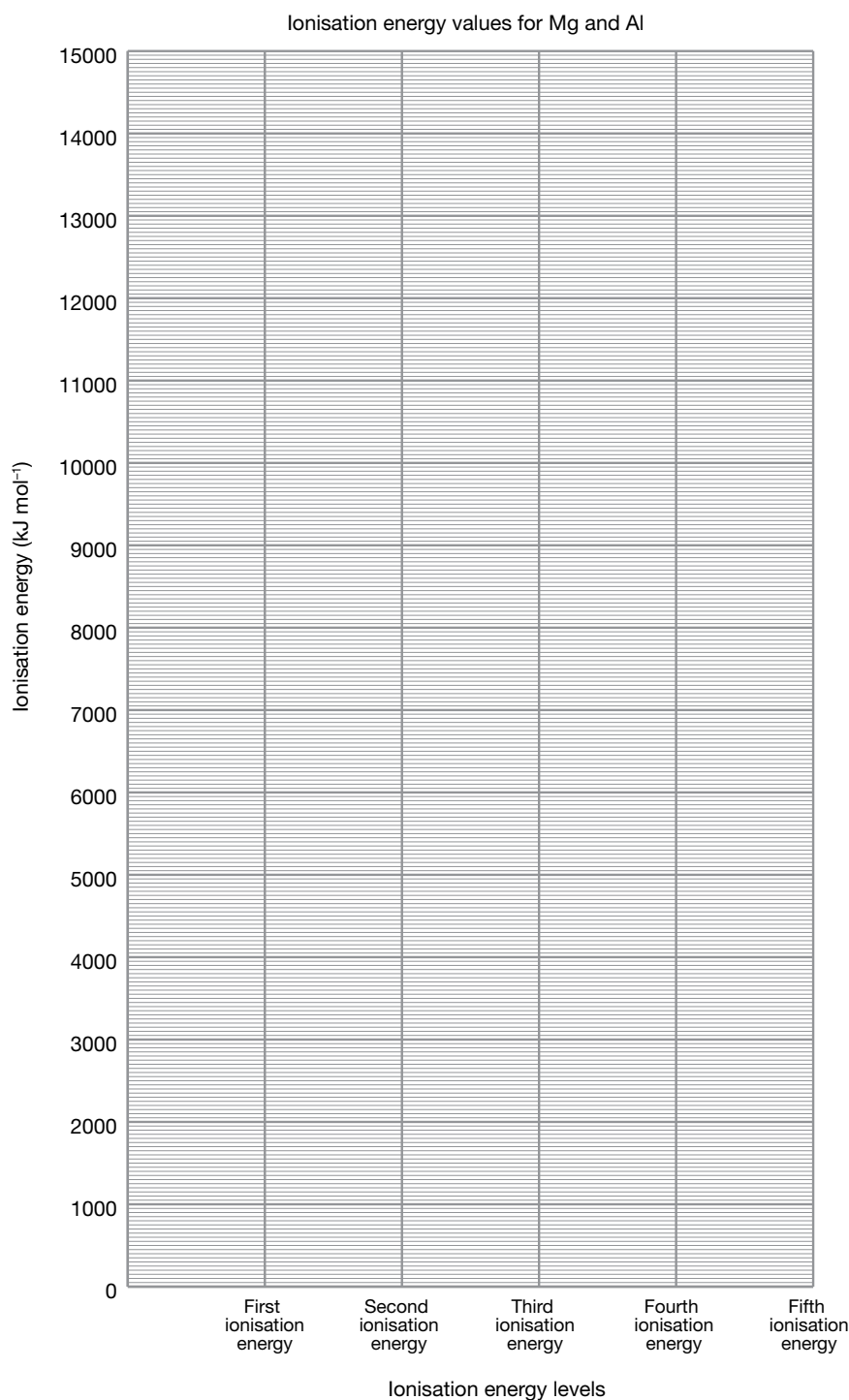
- (a) Using this information, graph first and second ionisation energy against atomic number.
- (b) Describe and explain two trends observed in ionisation energy down group 1.



12.1.2.2 The table shows the first five ionisation energies for the elements aluminium and magnesium.

Ionisation energy (kJ mol^{-1})	Magnesium	Aluminium
First ionisation energy	738	578
Second ionisation energy	1450	1820
Third ionisation energy	7730	2750
Fourth ionisation energy	10 500	11 600
Fifth ionisation energy	13 600	14 800

- (a) Using this information, graph ionisation energy values of the first to fifth ionisation energy levels, for both magnesium and aluminium.





(b) Describe and explain any trend in ionisation energy common to both graphs.

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(c) Describe and explain any differences between observed trends in ionisation energy for the following.

(i) Magnesium.

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(i) Aluminium.

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12.1.2.3 Outline the relationship between electron configuration of elements and ionisation energy.

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12.1.3 State the relative energies of s, p, d and f orbitals in a single energy level. © IBO 2007

12.1.3.1

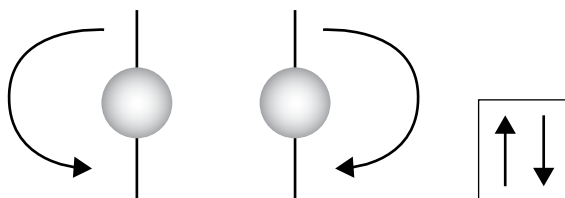
(a) Describe what is meant by an orbital.

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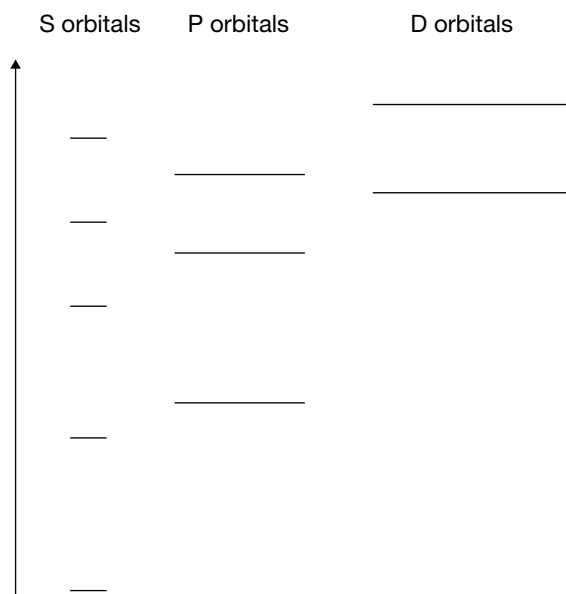
(b) The diagram below is a model designed to help students understand electron arrangements in atoms. What does it represent?



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12.1.3.2 Label the diagram to illustrate the relative electron energy levels in atoms.



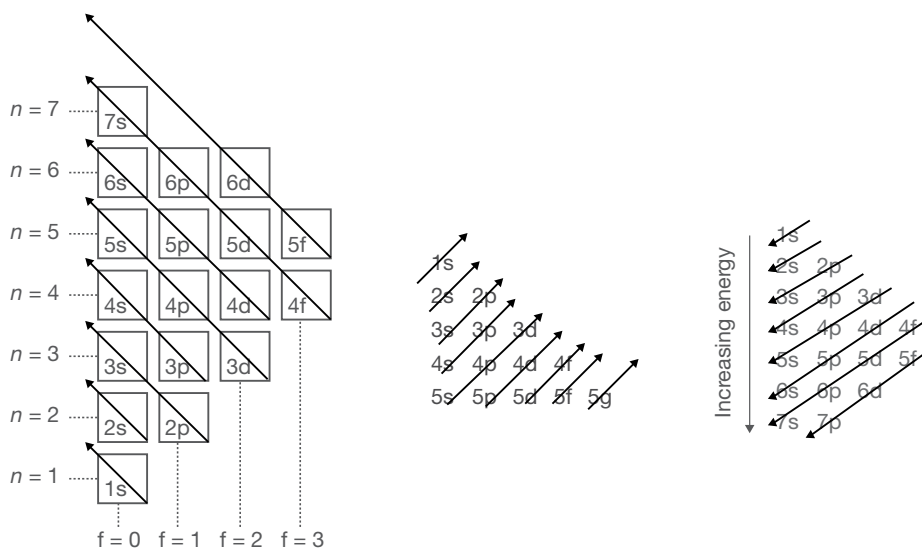
12.1.4 State the maximum number of orbitals in a given energy level. © IBO 2007

12.1.4.1 Each atom contains electrons orbiting the nucleus and arranged into energy levels and sublevels.

(a) Complete the table to describe the sublevels and orbitals of each atomic energy level and the electron capacity of each.

Energy level	Types of sublevels	Number of orbitals in each sublevel	Maximum number of electrons
1	s	1	2
2	s p	Total 4	
3		1 3 5 Total 9	
4		1 3 5 7 Total	32
n		n^2	

(b) The following diagrams all claim to show the order in which sublevels are filled – but they all have arrows pointing in different directions. Comment on these diagrams.



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- (c) Explain why the first row d-block element chromium is $[\text{Ar}] 3d^5 4s^1$ rather than $[\text{Ar}] 3d^4 4s^2$.

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12.1.5 Draw the shape of an s orbital and the shapes of the p_x , p_y and p_z orbitals. © IBO 2007

12.1.5.1 Different types of orbitals are distinguished by their shapes. Use diagrams to show the shapes of the following.

- (a) An s orbital.

- (b) Three types of p orbitals.

12.1.6 Apply the Aufbau principle, Hund's rule and the Pauli exclusion principle to write electron configurations for atoms and ions up to $Z = 54$. © IBO 2007

12.1.6.1 Outline what is meant by each of the following.

- (a) The Aufbau principle.

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- (b) Hund's rule.

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- (c) The Pauli exclusion principle.

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(d) The Heisenberg uncertainty principle.

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(e) The dual nature of electrons.

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12.1.6.2

(a) Complete the table below to show electron arrangement in the first 12 elements of the periodic table.

Element	Symbol	Electronic configuration	Orbitals with arrows to indicate number of electrons and their spin
Hydrogen		1.	$\uparrow$ 1s
Helium		2.	
	Li	2.1	
	Be		
	B		
	C		
Nitrogen			
Oxygen			
Fluorine			
	Ne		
	Na		
	Mg		

- (b) Complete the table to show the arrangement of electrons in the elements listed. The first one is done for you.

Atomic number	Name and symbol of element	Electrons in orbitals of level 1 (K)	Electrons in orbitals of level 2 (L)	Electrons in orbitals of level 3 (M)	Electrons in orbitals of level 4 (N)
		s	s p	s p d	s p d f
13	Aluminium (Al)	2	2 6	2 1	
18	Argon (Ar)				
19				2 6 -	
	Calcium (Ca)				
		2	2 6	2 6 1	2
26	Iron (Fe)				
	Nickel (Ni)				
			2 6	2 6 10	1
30					
31	(Ga)				
34	Selenium (Se)	2			2 4

12.1.6.3

(a) In an electron configuration, what would be meant by $2s^1$?

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(b) Write electron configurations, including orbital notation for the following elements.

(i) Silicon.

(ii) Phosphorus.

(iii) Chlorine.

(iv) Argon.

(v) Bromine.

(c) Identify the elements with the following notations.

(i) $1s^2 2s^2 2p^6 3s^2 3p^4$

(ii) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^2$

(iii) $1s^2 2s^2 2p^6 3s^1$

(iv) $1s^2 2s^2 2p^6 3s^2 3p^6 4s^1$

(v) $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^3$

(d) Different species with the same electronic structure are described as isoelectric. Identify three isoelectric species with the electronic structure $1s^2 2s^2 2p^6 3s^2 3p^6$.

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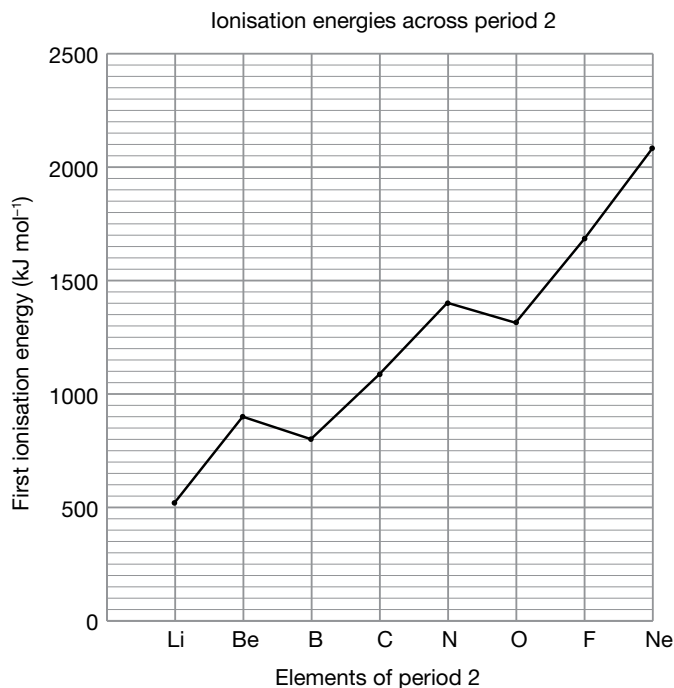
Answers



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AHL 12 Atomic Structure

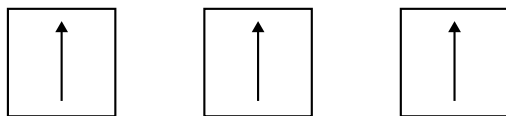
12.1.1.1 (a)



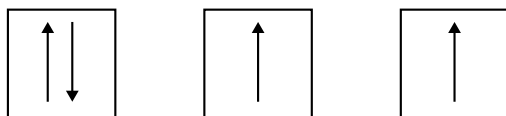
- (b) Ionisation energy generally increases across period 2 of the periodic table, from 520 kJ for lithium to 2081 kJ for neon. There is a slight drop from Be to B and from N to O.
- (c)
- The charge on the nucleus. A higher positive nuclear charge means a stronger attraction to negatively charged electrons and thus a higher ionisation energy. As you move across a period, the positive nuclear charge increases (number of protons increase) so the ionisation energy increases.
 - The effective nuclear charge. Electrons in inner shells have a shielding effect: each electron in a filled inner shell will cancel one unit of nuclear charge, e.g. beryllium has 4 protons, so that would produce a charge of +4 on the nucleus. However, the innermost shell has 2 electrons, so this reduces the effective nuclear charge (ENC) to +2.
 - The repulsion caused by the presence of other electrons within the shell from which the electron is being removed. When an electron is removed from a shell, there is less electron-electron repulsion and the remaining electrons move closer to the positive nucleus, making them harder to remove so their ionisation energy is greater.
- (d) Ionisation energy decreases from beryllium to boron and from nitrogen to oxygen. This is because:

Be to B – The electron lost from beryllium comes from an s subshell. The electron lost from boron comes from a p subshell which is at a higher energy than the s subshell and thus counteracts the effect of the increase in nuclear charge (B has 5 protons, Be has 4). The result is that the ionisation energy actually drops from 900 to 799 kJ/mol.

N to O – In N, each of the p orbitals is filled with 1 electron.

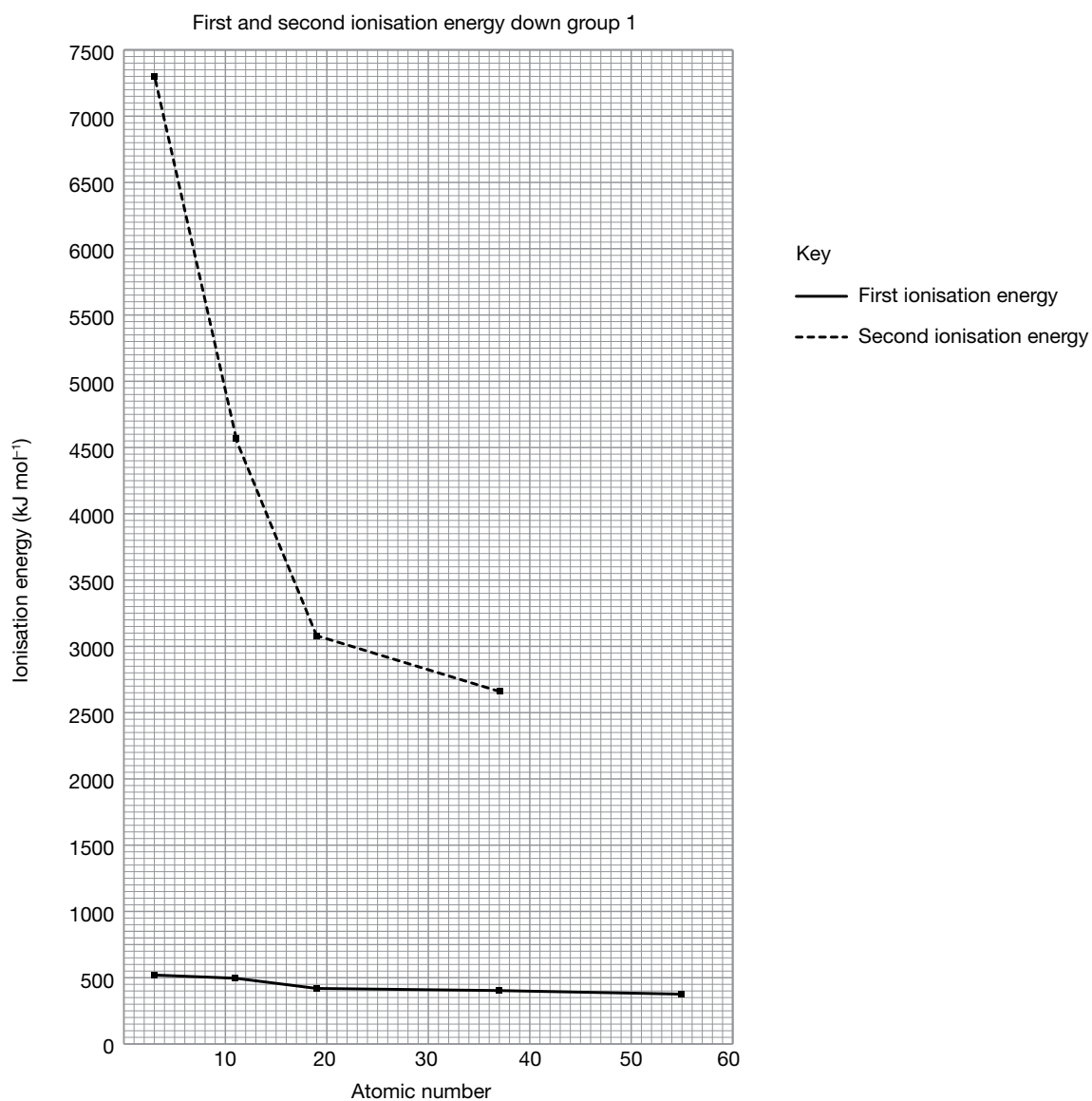


In O two electrons share an orbital.



Although there is an increase in nuclear charge (7 to 8 protons), this is more than counteracted by the greater electron-electron repulsion between the 2 electrons that share the same orbital in oxygen atoms and the ionisation energy drops from 1400 to 1310 kJ/mol.

12.1.2.1 (a)



(b) Trend: As you move down group 1, both first and second ionisation energies decrease.

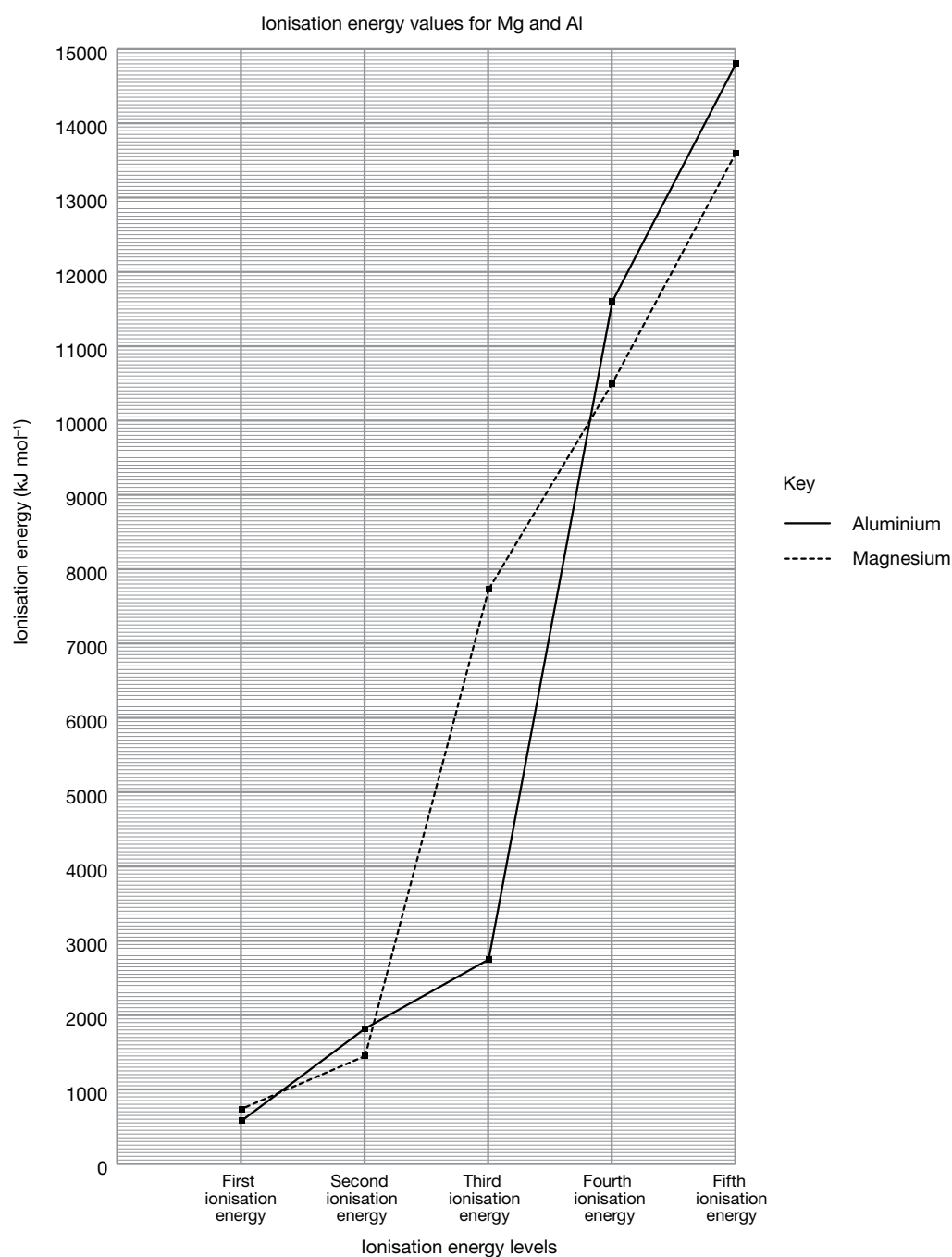
Explanation: As you move down the group more electron shells are being added. So the outermost electrons are further from the nucleus and thus they are attracted less strongly and need less energy to remove them from the atom.

Trend: Second ionisation energies are higher than first ionisation energies for all elements.

Explanation: These elements are all in group 1 so they each have one electron in their outer shell. When a second electron is removed, it comes from a full, stable shell, so more energy is needed to remove it.

12.1.2.2

(a)



- (b) In both graphs, there is an increase from the first to the fifth ionisation energy; as more electrons are removed from an atom, more energy is needed to remove the next electron.
- (c) First to fifth ionisation energy in magnesium and aluminium does not increase at a uniform rate.
- (i) Magnesium: There is a sudden increase from the second to the third ionisation. This suggests that magnesium has 2 electrons in its outermost valence shell, the third to fifth electrons coming from a second electron shell which is initially stable. The first 2 electrons (the valence electrons) have the same ENC (effective nuclear charge) of +2. The first one to be removed is also being repelled by the other one. Once one is removed, there is nothing to repel the second valence electron so it is attracted closer to the nucleus, and it has a higher ionisation energy (more energy is needed to remove it).
- The third, fourth and fifth electrons are being removed from the second energy level, so they all have an ENC of +10. As the first of these is removed, there are 5 other electrons in the p subshell which assist by repelling it (like charges repel). With each further removal there are fewer electrons in the same shell, so there is less helpful repulsion and more energy is needed to remove the electron – the ionisation energy is greater.

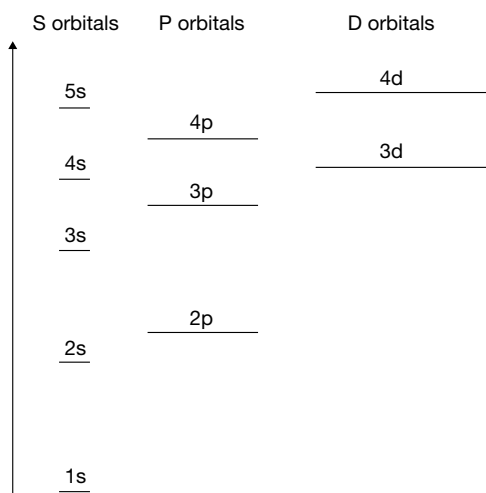
- (ii) Aluminium: The sudden increase in ionisation for aluminium comes after the third electron is removed. This suggests that aluminium has 3 electrons in its outer valence shell and the next one is drawn from an inner, stable shell. The first 3 electrons (the valence electrons) have the same ENC (+3). The first one to be removed is also being repelled by the other two. Once two are removed, there is nothing to repel the third valence electron and it is attracted closer to the nucleus, so it has higher ionisation energy (more energy is needed to remove it).

The fourth and fifth electrons are being removed from the second energy level, so they all have an ENC of +10. As the first of these is removed, there are 5 electrons left in the p subshell and they help by repelling it (like charges repel). With each further removal there are fewer electrons in the same shell, so there is less helpful repulsion and more energy is needed to remove these remaining electrons – their ionisation energy is greater.

12.1.2.3 Ionisation energy values allow us to determine the arrangement of electrons into levels and the electron configuration reflects this. For example, the electron configuration of magnesium is 2,8,2. We know this because measured ionisation energy values have determined that the electrons are arranged in levels – 2 in the first level, 8 in the second level and 2 in the third level.

- 12.1.3.1** (a) An orbital is a region of space, around an atom, which is occupied by 1 or 2 electrons in a section of an energy sublevel. The electrons have opposite spins. Within that region the electron has a particular energy. Examples of orbitals are the 1s, 2s and $3p_x$ orbitals.
- (b) The diagram represents 2 electrons in an orbital spinning in opposite directions. It shows that these can be represented as 2 arrows, one pointing up and the other pointing down to represent opposite spins.

12.1.3.2 Labels lines in order as:



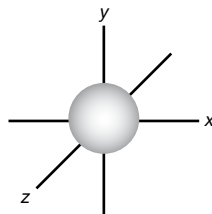
12.1.4.1 (a)

Energy level	Types of sublevels	Number of orbitals in each sublevel	Maximum number of electrons
1	s	1	2
2	s	1	8
	p	3	
	Total	4	
3	s	1	18
	p	3	
	d	5	
	Total	9	
4	s	1	32
	p	3	
	d	5	
	f	7	
	Total	16	
n	n types	n^2	$2n^2$

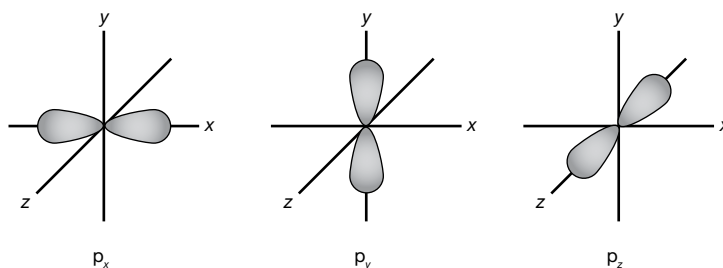
- (b) They all show the same thing, they are just arranged differently. They all show the order of filling sublevels is: 1s, 2s, 2p, 3s, 3p, 4s, 3d, 4p, 5s, 4d, 5p, 6s, 4f, 5d, 6p, 7s, 5f, 6d, 7p.
- (c) The ion with the most d orbitals is more stable. $3d^5 4s^1$ indicates there is 1 electron in each of the 5 d orbitals, leaving 1 electron in the 4s orbital. This arrangement is more stable than leaving a d orbital empty and having 2 electrons in the 4s orbital ($3d^4 4s^2$).

Note: This also applies to copper.

12.1.5.1 (a) The s orbitals are round in shape.



(b) The p orbitals are shaped like a figure of eight. The three types vary with their orientation.



2p atomic orbitals

- 12.1.6.1 (a) The Aufbau (building up) principle states that electrons in atoms always take the lowest energy configuration possible. They fill one subshell completely before starting to fill the subshell with the next highest energy level.
- (b) Hund's rule states that electrons adopt the most stable configuration.
- (c) The Pauli exclusion principle states that no two electrons in an atom can be in the same place at the same time. (No two electrons in a given atom can have the same four quantum numbers.)
- (d) The Heisenberg uncertainty principle states that it is impossible to determine precisely where an electron is in an atom or where it is going next. (Thus you cannot plot an orbit for an electron.)
- (e) Electrons are considered to have a dual nature because they have properties of particles (as in the Bohr theory of the atom) and they also have properties of waves (e.g. they are diffracted when they pass through a small slit – as are water and light waves).

12.1.6.2

Element	Symbol	Electronic configuration	Orbitals with arrows to indicate number of electrons and their spin
Hydrogen	H	1.	$\uparrow$ 1s
Helium	He	2.	$\uparrow\downarrow$ 1s
Lithium	Li	2.1	$\uparrow\downarrow$ $\uparrow$ 1s 2s
Beryllium	Be	2.2	$\uparrow\downarrow$ $\uparrow\downarrow$ 1s 2s
Boron	B	2.3	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow$ — — 1s 2s 2p
Carbon	C	2.4	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow$ $\uparrow$ — 1s 2s 2p
Nitrogen	N	2.5	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow$ $\uparrow$ $\uparrow$ 1s 2s 2p
Oxygen	O	2.6	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow$ $\uparrow$ 1s 2s 2p
Fluorine	F	2.7	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow$ 1s 2s 2p
Neon	Ne	2.8	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ 1s 2s 2p
Sodium	Na	2.8.1	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow$ — — 1s 2s 2p 3s 3p
Magnesium	Mg	2.8.2	$\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ $\uparrow\downarrow$ — — 1s 2s 2p 3s 3p

(b)

Atomic number	Name and symbol of element	Electrons in orbitals of level 1 (K)	Electrons in orbitals of level 2 (L)	Electrons in orbitals of level 3 (M)	Electrons in orbitals of level 4 (N)
		s	s p	s p d	s p d f
13	Aluminium (Al)	2	2 6	2 1	
18	Argon (Ar)	2	2 6	2 6	
19	Potassium (K)	2	2 6	2 6 –	1
20	Calcium (Ca)	2	2 6	2 6 –	2
21	Scandium (Sc)	2	2 6	2 6 1	2
26	Iron (Fe)	2	2 6	2 6 6	2
28	Nickel (Ni)	2	2 6	2 6 8	2
29	Copper (Cu)	2	2 6	2 6 10	1
30	Zinc (Zn)	2	2 6	2 6 10	2
31	Gallium (Ga)	2	2 6	2 6 10	2 1
34	Selenium (Se)	2	2 6	2 6 10	2 4

12.1.6.3

- (a) The 2 is the energy level (the principal quantum number), s is the type of subshell, and the superscript 1 tells us the number of electrons in that subshell.
- (b) (i) Silicon – $1s^2 2s^2 2p^6 3s^2 3p^2$
(ii) Phosphorus – $1s^2 2s^2 2p^6 3s^2 3p^3$
(iii) Chlorine – $1s^2 2s^2 2p^6 3s^2 3p^5$
(iv) Argon – $1s^2 2s^2 2p^6 3s^2 3p^6$
(vi) Bromine – $1s^2 2s^2 2p^6 3s^2 3p^6 3d^{10} 4s^2 4p^5$
- (c) (i) Sulfur
(ii) Calcium
(iii) Sodium
(iv) Potassium
(v) Arsenic
- (d) Answer could include: atoms of argon (Ar) and the following ions: chloride ion (Cl^-), sulfide ion (S^{2-}), potassium ion (K^+) and calcium ion (Ca^{2+}).